

T08D06 - Dissociation constants of Acids and Bases

Part I: strong and weak acids and bases:

Strong Acids	Weak Acids	Weak Bases	Strong Bases
HClO ₄	HF	NH ₃	LiOH
HI	H ₂ CO ₃	N ₂ H ₄	NaOH
HBr	H ₂ S	(CH ₃) ₂ NH	KOH
HCl	H ₃ BO ₃	C ₆ H ₅ N	RbOH
H ₂ SO ₄	HC ₂ H ₃ O ₂	CH ₃ NH ₂	CsOH
HNO ₃	HC ₇ H ₅ O ₂	NaHCO ₃	Ca(OH) ₂
	C ₆ H ₄ (OH)CO ₂ CH ₃		Sr(OH) ₂
			Ba(OH) ₂

Part II: Definitions:

Dissociation Process by which the action of a solvent or a change in physical condition, as in pressure or temperature, causes a molecule to split into simpler groups of atoms, single atoms, or ions.

Ionization Process by which a neutral compound is split into charged particles by action when dissolved in liquid water.

Equilibrium When reactants and products are in a constant stable ratio. The forward & reverse reactions occur @ the same rate when a system is in equilibrium.

Completion

vs.

Equilibrium

→ All products (if no excess)

⇌ starting materials left unreacted in part.

Part III: How to calculate the pH, pOH, [H⁺], [OH⁻] of acids and bases – write the equilibrium constant to figure it out:

Strong Acids and Bases: $\text{HA} + \text{H}_2\text{O} \rightarrow \text{H}^+ + \text{H}_3\text{O}^+$ (for acids)

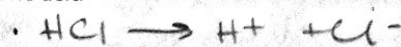
Weak Acids and Bases: $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{H}_3\text{O}^+$ (for acids)

Notice, the only difference is in the arrow for completion or equilibrium!

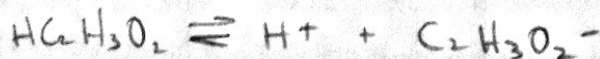
Part IV: Ionization reactions.

Write the ionization reactions for each of the following reactions [include water in the reaction. Include the states for each-aq or l.

a. hydrochloric acid



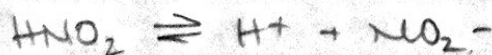
b. acetic acid



c. ammonium ion



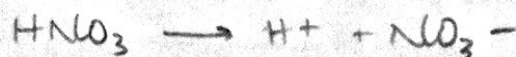
d. nitrous acid



e. sulfuric acid



f. nitric acid



Yields
↓
Strong acids →
Weak acids ⇌
↑
Equilibrium

Part V: Write the equilibrium expression for the following letters from part IV:

b. $K_{eq} = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$

c. $K_a = \frac{[H^+][NH_3]}{[NH_4^+]}$

d. $K_a = \frac{[H^+][NO_2^-]}{[HNO_2]}$

f. $K_a = \text{STRONG ACID} = \text{CANNOT WRITE}$

Part VI: Equilibrium Constant

Calculate the pH, pOH, $[H^+]$, and $[OH^-]$ concentration of each of the following:

a. 1.0 M solution of HNO_3 (strong)

$$-\log[1.0] = 0 = pH \quad [H^+] = 1 \times 10^0$$

$$14 = pOH \quad [OH^-] = 1 \times 10^{-14}$$

b. 0.65 M solution of NaOH (strong)

$$-\log[0.65] = 0.187 = pOH \quad [OH^-] = 6.5 \times 10^{-1}$$

$$13.81 = pH \quad [H^+] = 1.54 \times 10^{-14}$$

c. 13.45 g of H_2SO_4 placed in water to make 325 mL of solution. [Diprotic acid but only the first ionization step should be used.] (strong)

$$13.45 \text{ g } H_2SO_4 \times \frac{1 \text{ mol } H_2SO_4}{98.09 \text{ g } H_2SO_4} = 0.137 \text{ mol } H_2SO_4$$

$$M = \frac{0.137 \text{ mol}}{0.325 \text{ L}} = 0.422 \text{ M } H_2SO_4$$

$$-\log[0.442] = 0.375 = pH \quad [H^+] = 4.42 \times 10^{-1}$$

$$13.61 = pOH \quad [OH^-] = 2.38 \times 10^{-14}$$

d. 36.5 g of $Ba(OH)_2$ placed in water to make 965 mL of solution. (strong)

$$36.5 \text{ g } Ba(OH)_2 \times \frac{1 \text{ mol } Ba(OH)_2}{171.35 \text{ g } Ba(OH)_2} = 0.213 \text{ mol } Ba(OH)_2$$

$$\frac{0.213 \text{ mol } Ba(OH)_2 \times 2 \text{ mol } OH^-}{0.965 \text{ L}} \times \frac{1 \text{ mol } Ba(OH)_2}{1 \text{ mol } Ba(OH)_2} = 0.440 \text{ M } OH^-$$

$$-\log[0.440] = 0.36 = pOH \quad [OH^-] = 4.40 \times 10^{-1}$$

$$13.64 = pH \quad [H^+] = 2.27 \times 10^{-14}$$

e. a 1.5 M solution of KOH (strong)

$$-\log[1.5] = pOH = 0$$

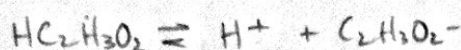
$$pH = 14$$

$$[OH^-] = 1.5 \times 10^0$$

$$[H^+] = 6.67 \times 10^{-15}$$

- f. a 0.50 M solution of a acetic acid that is a weak acid ($K_a = 1.8 \times 10^{-5}$)

weak



I	0.5M	0	0
C	-x	+x	+x
E	0.5-x	+x	+x

$$1.8 \times 10^{-5} = K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{[x][x]}{[0.5-x]} = \frac{x^2}{0.5-x}$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$0 = x^2 + 1.5 \times 10^{-5}x - 9.0 \times 10^{-6}$$

$$a = 1$$

$$b = 1.5 \times 10^{-5}$$

$$c = -9.0 \times 10^{-6}$$

* Assumption: If K_a is 1000x smaller than C , then is no need for the quadratic function.

$$x = \frac{-1.5 \times 10^{-5} \pm \sqrt{(1.5 \times 10^{-5})^2 - 4(1)(-9.0 \times 10^{-6})}}{2(1)}$$

$$\frac{x^2}{0.5-x} = \frac{x^2}{0.5} = 1.8 \times 10^{-5}$$

$$x = 0.02999$$

$$\approx 0.0300$$

$$\boxed{\begin{array}{l} x = 0.003 \text{ M} = [\text{H}^+] \\ 3.33 \times 10^{-12} = [\text{H}^+] \\ 2.52 = \text{pH} \\ 11.48 = \text{pOH} \end{array}}$$

Since we used the assumption we should be close, in result our answers were just fine by the assumption

- g. Acetic acid has a $K_a = 1.78 \times 10^{-5}$. Determine the pH of a 0.100 M solution.

$$1.78 \times 10^{-5} = K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{x^2}{0.100-x} \approx \frac{x^2}{0.100} \quad x = 1.33 \times 10^{-3} = [\text{H}^+]$$

$$-\log[\text{H}^+] = \text{pH} = 2.87$$

- h. Calculate the hydronium ion ($[\text{H}_3\text{O}^+]$) concentration of a 0.25 M solution of benzoic acid,

$$\text{HC}_7\text{H}_5\text{O}_2, K_a = 6.46 \times 10^{-5}$$

$$6.46 \times 10^{-5} = K_a = \frac{[\text{H}^+][\text{C}_7\text{H}_5\text{O}_2^-]}{[\text{HC}_7\text{H}_5\text{O}_2]} = \frac{x^2}{0.25-x} \approx \frac{x^2}{0.25}$$

$$x = 4.02 \times 10^{-3} = [\text{H}^+] = [\text{H}_3\text{O}^+]$$

- i. Hydrocyanic acid, HCN, has a K_a of 4.93×10^{-10} . Find the pOH and $[\text{H}_3\text{O}^+]$ of a 0.100 M solution of hydrocyanic acid.

$$4.93 \times 10^{-10} = K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]} = \frac{x^2}{0.100-x} \approx \frac{x^2}{0.100}$$

$$x = 7.02 \times 10^{-6} = [\text{H}^+] = [\text{H}_3\text{O}^+]$$

$$-\log[\text{H}^+] = \text{pH} = 5.15$$

$$\text{pOH} = 8.85$$

- j. The weak base analine, $\text{C}_6\text{H}_5\text{NH}_2$, has a K_b of 4.26×10^{-10} . Find the pH, pOH and $[\text{H}_3\text{O}^+]$ of a 0.100 M solution of analine.

$$4.26 \times 10^{-10} = K_b = \frac{[\text{C}_6\text{H}_5\text{NH}_2^+][\text{OH}^-]}{[\text{C}_6\text{H}_5\text{NH}_2]} = \frac{x^2}{0.100-x} \approx \frac{x^2}{0.100}$$

$$x = 6.53 \times 10^{-6} = [\text{OH}^-] \quad \text{pOH} = 5.19$$

$$1.53 \times 10^{-9} = [\text{H}^+] \quad \text{pH} = 8.81$$

- k. Calculate the pH of a 0.75 M solution of a methylamine, CH_3NH_2 whose $K_b = 4.54 \times 10^{-4}$.

$$4.54 \times 10^{-4} = K_b = \frac{[\text{CH}_3\text{NH}_2^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = \frac{x^2}{0.75-x} \approx \frac{x^2}{0.75}$$

$$x = 1.85 \times 10^{-2} = [\text{OH}^-] \quad \text{pOH} = 1.73$$

$$5.41 \times 10^{-13} = [\text{H}^+] \quad \text{pH} = 2.27$$

- l. Calculate the pH of a 0.155 M ammonia solution. $K_b = 1.77 \times 10^{-5}$

$$1.77 \times 10^{-5} = K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{x^2}{0.155-x} \approx \frac{x^2}{0.155}$$

$$x = 1.66 \times 10^{-3} = [\text{OH}^-] \quad \text{pOH} = 2.78$$

$$6.04 \times 10^{-12} = [\text{H}^+] \quad \text{pH} = 11.22$$

- m. Chloroacetic acid, ClCH_2COOH , has a K_a of 1.40×10^{-3} . Find the pH and $[\text{OH}^-]$ of a 0.30 M solution of chloroacetic acid.

$$1.40 \times 10^{-3} = K_a = \frac{[\text{ClCH}_2\text{COO}^-][\text{H}^+]}{[\text{ClCH}_2\text{COOH}]} = \frac{x^2}{0.30-x} \xrightarrow{\text{Assumption}} \frac{x^2}{0.30}$$

$$x = \sqrt{2.05 \times 10^{-2}} = [\text{H}^+] \quad \text{pH} = 1.69$$
$$4.88 \times 10^{-13} = [\text{OH}^-] \quad \text{pOH} = 12.31$$